(1) (a) acidic
(b) basic
(c) neutral
(d) acidic
(e) neutral ($K_a = K_b$)

(2) $HNO_{2(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + NO_2^- \uparrow$

$\xleftarrow{\text{shifts left}}$

HNO_2 will dissociate less in $NaNO_2$ because the presence of more NO_2^- in solution will shift HNO_2's equilibrium left which will result in less ionization.

(3) Acid added: $NH_{3(aq)} + H^+_{(aq)} \longrightarrow NH_4^+{}_{(aq)}$
Base added: $NH_4^+{}_{(aq)} + OH^-_{(aq)} \longrightarrow NH_{3(aq)} + H_2O_{(l)}$

(4) (a) $pH = pK_a + \log \frac{[BrO^-] \leftarrow \text{base}}{[HBrO] \leftarrow \text{acid}}$ $K_{a,HBrO} = 2.8 \times 10^{-9}$

$$pH = -\log(2.8\times10^{-9}) + \log\left(\frac{0.18}{0.15}\right) = 8.63$$

(b) $HBrO_{(aq)} + NaOH_{(aq)} \longrightarrow NaBrO_{(aq)} + H_2O_{(l)}$

	$HBrO_{(aq)}$	$NaOH_{(aq)}$	$NaBrO_{(aq)}$	$H_2O_{(l)}$
Before Addition	0.15 mol	0	0.18 mol	
Addition	—	0.01 mol	—	
After Addition	0.14 mol	0	0.19 mol	

$$pH = -\log(2.8\times10^{-9}) + \log\left(\frac{0.19 \leftarrow \text{base}}{0.14 \leftarrow \text{acid}}\right) = \underline{8.69}$$

(c) $NaBrO_{(aq)} + HCl_{(aq)} \longrightarrow HBrO_{(aq)} + NaCl_{(aq)}$

	$NaBrO_{(aq)}$	$HCl_{(aq)}$	$HBrO_{(aq)}$	$NaCl_{(aq)}$
Before Addition	0.18 mol	0	0.15 mol	
Addition	—	0.01 mol	—	
After Addition	0.17 mol	0	0.16 mol	

$$pH = -\log(2.8\times10^{-9}) + \log\left(\frac{0.17}{0.16}\right) = 8.58$$

base (0.17), acid (0.16)

(5) $pH = pKa + \log\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}$ ← base / ← acid 1 $Ka, HC_2H_3O_2 = 1.8\times10^{-5}$

$5.00 = 4.74 + \log\frac{[C_2H_3O_2^-]}{0.150}$ 1

$10^{0.26} = \frac{[C_2H_3O_2^-]}{0.150} \Rightarrow [NaC_2H_3O_2] = (0.150)(10^{0.26}) = \underline{0.27\,M}$ 1

mol $NaC_2H_3O_2 = (0.27M)(1.00L) = 0.27$ mol $NaC_2H_3O_2$ 1

$0.27 \text{ mol } NaC_2H_3O_2 \left(\frac{82.04 \text{ g } NaC_2H_3O_2}{1 \text{ mol } NaC_2H_3O_2}\right) = \underline{22 \text{ g } NaC_2H_3O_2}$ 1

(7) Equivalence Point: $C_aV_a = C_bV_b \Rightarrow V_a = \frac{C_bV_b}{C_a} = \frac{(0.240M)(25.0mL)}{0.200\,M}$

Strong base – Strong acid titration

$V_a = 30.0$ mL Equivalence Point

(a) 0.0 mL

$[OH^-] = [NaOH] = 0.240$ M

$pOH = -\log(0.240) = 0.620$

$pH = 14.00 - 0.62 = \underline{13.38}$

(b) 15.0 mL

Excess NaOH

	$NaOH_{(aq)}$ +	$HNO_{3(aq)}$ →	$NaNO_{3(aq)}$ +	$H_2O_{(l)}$
BA	6.00 mmol	—	—	
A	—	3.00 mmol	—	
AA	3.00 mmol	0	3.00 mmol	

← neutral salt.

$[OH^-] = \frac{3.00 \text{ mmol}}{40.0 \text{ mL}} = 0.0750$ M

$pOH = -\log(0.0750) = 1.125 \Rightarrow pH = 14.00 - 1.125 = \underline{12.88}$

Excess NaOH

(c) 25.0 mL

	$NaOH_{(aq)}$ +	$HNO_{3\,(aq)}$ $\longrightarrow$	$NaNO_{3\,(aq)}$ +	$H_2O_{(l)}$
BA	6.00 mmol	—	—	
A	—	5.00 mmol	—	
AA	1.00 mmol	0	5.00 mmol	

neutral salt

$$[OH^-] = \frac{1.00 \text{ mmol}}{50.0 \text{ mL}} = 0.0200\,M$$

$$pOH = -\log(0.0200) = 1.699 \Rightarrow pH = 14.00 - 1.699 = \underline{\underline{12.30}}$$

Equivalence point

(d) 30.0 mL

	$NaOH_{(aq)}$ +	$HNO_{3\,(aq)}$ $\longrightarrow$	$NaNO_{3\,(aq)}$ +	$H_2O_{(l)}$
BA	6.00 mmol	—	—	
A	—	6.00 mmol	—	
AA	0	0	6.00 mmol	

Neutral salt

At equivalence point of a strong base - strong acid titration, a neutral salt is present in solution (no excess of strong acid or base present) so pH is 7.00.

Excess HNO_3

(e) 45.0 mL

	$NaOH_{(aq)}$ +	$HNO_{3\,(aq)}$ $\longrightarrow$	$NaNO_{3\,(aq)}$ +	$H_2O_{(l)}$
BA	6.00 mmol	—	—	
A	—	9.00 mmol	—	
AA	0	3.00 mmol	6.00 mmol	

neutral salt

$$[H_3O^+] = \frac{3.00 \text{ mmol}}{70.0 \text{ mL}} = 0.0429\,M$$

$$pH = -\log(0.0429) = \underline{\underline{1.368}}$$

⑥ (a) weak acid / strong base titration pH is above 7.
(b) strong acid / strong base titration pH is 7.
(c) weak base / strong acid titration pH is below 7.

⑧ Equivalence Volume $C_aV_a = C_bV_b \Rightarrow V_a = \frac{C_bV_b}{C_a}$

$= \frac{(25.0\text{ mL})(0.240\text{ M})}{0.200\text{ M}}$

$V_a = 30.0\text{ mL}$

weak base solution

(a) 0.0 mL HCl $[OH^-] = \sqrt{K_b \times [CH_3NH_2]}$

$= \sqrt{(4.4\times10^{-4})(0.240)} = 1.0\times10^{-2}\text{ M}$

$pOH = -\log[OH^-] = -\log(1.0\times10^{-2}) = 1.99$

$pH = 14.00 - 1.99 = \underline{12.01}$

(b) 15.0 mL HCl Mid Point of titration $V_a = \frac{1}{2}V_e$

Buffer Region

K_a of $CH_3NH_3^+ \Rightarrow K_a = \frac{K_w}{K_b} = \frac{1.0\times10^{-14}}{4.4\times10^{-4}} = 2.3\times10^{-11}$

At midpoint, $pH = pK_a = -\log(2.3\times10^{-11}) = \underline{10.64}$

(c) 25.0 mL HCl

Buffer Region

	$CH_3NH_{2(aq)}$ +	$HCl_{(aq)}$ →	$CH_3NH_3Cl_{(aq)}$
BA	6.00 mmol	—	—
A	—	5.00 mmol	—
AA	1.00 mmol	0	5.00 mmol

$pH = pK_a + \log\frac{[base]}{[acid]} = -\log(2.3\times10^{-11}) + \log\frac{(1.00)}{(5.00)} = \underline{9.94}$

Equivalence Point (d) 30.0 mL HCl

	$CH_3NH_{2(aq)}$ +	$HCl_{(aq)}$ ⟶	$CH_3NH_3Cl_{(aq)}$
BA	6.00 mmol	—	—
A	—	6.00 mmol	—
AA	0	0	6.00 mmol

$[CH_3NH_3Cl] = \frac{6.00\ mmol}{55.0\ mL} = 0.109\ M$

$[H_3O^+] = \sqrt{K_a \times [CH_3NH_3Cl]} = \sqrt{(2.3\times10^{-11})(0.109)} = 1.6\times10^{-6}\ M$

$pH = -\log[H_3O^+] = -\log(1.6\times10^{-6}) = \underline{\underline{5.80}}$

Excess HCl (e) 45.0 mL HCl

	$CH_3NH_{2(aq)}$ +	$HCl_{(aq)}$ ⟶	$CH_3NH_3Cl_{(aq)}$
BA	6.00 mmol	—	—
A	—	9.00 mmol	—
AA	0	3.00 mmol	6.00 mmol

$[H_3O^+] = [HCl] = \frac{3.00\ mmol}{70.0\ mL} = 0.0429\ M$

$pH = -\log[H_3O^+] = -\log(0.0429) = \underline{\underline{1.368}}$

③ Equivalence Point: $C_aV_a = C_bV_b \Rightarrow V_b = \frac{C_aV_a}{C_b} = \frac{(0.200\ M)(25.0\ mL)}{0.250\ M}$

weak acid - strong base titration

$V_{eq} = 20.0\ mL$ Equivalence point.

Weak Acid Solution (a) 0.0 mL $[H_3O^+] = \sqrt{K_a \times [HC_2H_3O_2]} = \sqrt{(1.8\times10^{-5})(0.200)} = 1.9\times10^{-3}\ M$

$pH = -\log(1.9\times10^{-3}) = \underline{\underline{2.72}}$

Buffer Region

(b) 10.0 mL $V_b = \frac{1}{2} V_{eq}$ Mid Point of titration

At midpoint $pH = pKa \Rightarrow pH = -\log(1.8\times10^{-5}) = \underline{4.74}$

Buffer Region

(c) 15.0 mL

	$HC_2H_3O_{2(aq)}$ +	$KOH_{(aq)}$ →	$KC_2H_3O_{2(aq)}$ +	$H_2O_{(l)}$
BA	5.00 mmol	—	—	
A	—	3.75 mmol	—	
AA	1.25 mmol	0	3.75 mmol	

$$pH = -\log(1.8\times10^{-5}) + \log\left(\frac{3.75}{1.25}\right) = \underline{5.22}$$

Equivalence Point

weak base solution

(d) 20.0 mL

	$HC_2H_3O_{2(aq)}$ +	$KOH_{(aq)}$ →	$KC_2H_3O_{2(aq)}$ +	$H_2O_{(l)}$
BA	5.00 mmol	—	—	
A	—	5.00 mmol	—	
AA	0	0	5.00 mmol	

→ Basic salt

$$[C_2H_3O_2^-] = \frac{5.00\text{ mmol}}{45.0\text{ mL}} = 0.111\text{ M} \qquad K_b = \frac{1.0\times10^{-14}}{1.8\times10^{-5}} = 5.6\times10^{-10}$$

$$[OH^-] = \sqrt{K_b \times [C_2H_3O_2^-]} = \sqrt{(0.111)(5.6\times10^{-10})} = 7.9\times10^{-6}\text{ M}$$

$$pOH = -\log(7.9\times10^{-6}) = 5.10$$

$$pH = 14.00 - 5.10 = \underline{8.90}$$

Excess KOH

(e) 30.0 mL

	$HC_2H_3O_{2(aq)}$ +	$KOH_{(aq)}$ ⟶	$KC_2H_3O_{2(aq)}$ +	$H_2O_{(l)}$
BA	5.00 mmol	—		
A	—	7.50 mmol	—	
AA	0	2.50 mmol	5.00 mmol	

↳ basic salt (a lot weaker than KOH)

$[OH^-] = \frac{2.50\text{ mmol}}{55.0\text{ mL}} = 0.0455\,M$

$pOH = -\log(0.0455) = 1.342$

$pH = 14.00 - 1.342 = \underline{\underline{12.66}}$

(9) Plot titration curve ⇒ pH vs. mL NaOH added

Plot attached

At midpoint ($V_{NaOH} = \frac{1}{2} V_{equivalence}$), $pH = pKa$

Equivalence Volume ≅ 22.4 mL ; midpoint volume = 11.2 mL

pH @ midpoint ≅ 4.50 = pka ⇒ $Ka = 10^{-pka} = 10^{-4.50} = \underline{\underline{3.2 \times 10^{-5}}}$

Titration of a monoprotic acid with a NaOH solution